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Q1. Two characteristics of ions that determine how they pack together to form crystals are the ion charges and their relative sizes. How quickly the crystal forms will also affect packing.

Q2. The main reason why ions pack together with different types of symmetry is the relative charge of the ions. In the sodium chloride crystal, both sodium and chloride ions each have a charge of 1. So, they will pack in a ratio of 1:1. If the crystal is made of calcium and chloride ions, however, the ratio must be 1:2, since the charge on a calcium ion is 2^+, and the charge on chloride ion is only 1^-. The difference in the ratio will have a major effect on the crystal shape.

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For full solutions to the practice problems, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

1. linear

2. tetrahedral (pyramidal)

3. trigonal planar

4. trigonal planar

5. tetrahedral (pyramidal)

6. tetrahedral (bent)

7. tetrahedral

8. tetrahedral

9. linear

10. trigonal planar

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Q3. The features that must be obtained from the Lewis structures are the types of bonding pairs (LP or BP, how many of each), and whether the bonds are single, double, or triple.

Q4. "Electron group" is defined as the electrons involved in a bond. It may be a single, double, or triple bond.

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For full solutions to the practice problems, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

11. tetrahedral, polar

12. trigonal planar, polar

13. tetrahedral (pyramidal), polar

14. research shows that depending on the bends, the vectors may or may not cancel

15. tetrahedral, polar

16. $NF_3(g)$ more polar than $NCl_3(\ell)$

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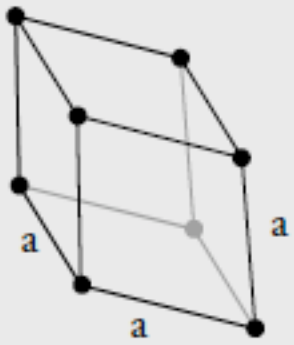
1. A unit cell is defined as the smallest set of ions in a crystal for which a pattern is repeated over and over. It is a three-dimensional structure that details how each of the faces will appear in the crystal. Sodium chloride, NaCl, has a cube-shaped unit cell (as shown in Figure 2.2(C) on page 48), whereas calcium carbonate, $CaCO_3$, has a rhombohedral shape (diamond), as shown below:



2. Cations (positive ions) are smaller than neutral atoms of the same element, whereas anions (negative ions) are larger than neutral atoms of the same element. The cations in any period of the periodic table are smaller than the anions of the same period.
3. The two factors that affect the shape of an ionic crystal are the relative size and charge of the ions in the crystal.
4. In an ionic crystal, the charged ions are simultaneously attracted to all of the oppositely charged ions surrounding them, which leads to the building of a crystal of potentially infinite size. In a molecule, however, the attractions, or bonds that hold the atoms in the molecule together, exist only between specific pairs of atoms.
5. VSEPR theory states that bonds and lone pairs spread out as much as possible in a molecule. By counting the number and type of bonding elements, the shape of the molecule can be predicted because it will maximize the distance between the bonding elements.
6. LP-LP (lone pair-lone pair) > LP-BP (bonding pair) > BP-BP. The order of repulsion will not affect the overall shape of the molecule but will affect the bond angle between the atoms in the molecule. For example, the theoretical angle between any two atoms in a tetrahedral shape is 109.5°. In a bent molecule, the bond angle would be 109.5° if lone pair repulsions

were not greater than either lone pair-bonding pair or bonding pair-bonding pair repulsions. However, in the bent molecule (water (H_2O)) the bond angle is 104.5° because of the greater repulsions between the lone pairs and the lone pair-bonding pairs than between the bonding pairs.

7. The type of bond, whether single, double, or triple does affect the molecular shape. A triple bond, although consisting of three bonding pairs, is only one bonding element. One bonding element leads to a linear shape, whereas three single bonds, or three bonding elements, would be either pyramidal (with one LP) or trigonal planar (no LP).

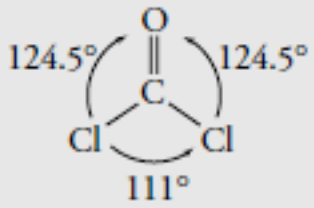
8.

Shape	Configuration leading to the shape	Types of bonds
Tetrahedral	4 bonding pairs, no lone pairs	4 single bonds
Trigonal Planar	3 bonding pairs 4 bonding pairs	3 single bonds, no lone pairs 1 double bond, 2 single bonds
Pyramidal	3 bonding pairs, one lone pair	3 single bonds
Bent	2 bonding pairs, 2 lone pairs 3 bonding pairs, 1 lone pair	2 single bonds 1 double bond, 1 single bond

Linear	1 bonding pair, 3 lone pairs 2 bonding pairs, 2 lone pairs 3 bonding pairs, 1 lone pair	1 single bond 1 double bond 1 triple bond

9. In a trigonal planar molecule, there are three electron groupings which are all bonding pairs. The ideal angle between the bonds is 120°.

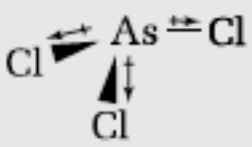
(a)



(b) The predicted angles of 120° are based on each of the repulsions being equal. Based on the observed bond angles, we could infer that the repulsions from a double bond are greater than the repulsions from a single bond, leading to the deviations observed.

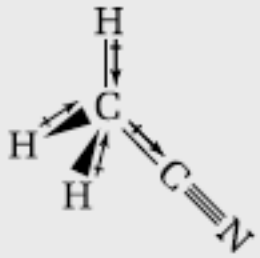
(c) The bond angles between the oxygen and hydrogen in methanal are 121°, compared to 124.5° observed in this molecule. The repulsion between the oxygen and chlorine atoms is greater than the repulsion between oxygen and hydrogen atoms. Since chlorine has three lone pairs, there should be more repulsion between oxygen and chlorine, as observed.

10. (a)



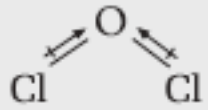
Trigonal pyramidal molecule: the top (arsenic end) of the pyramidal shape is positive, the base vertices (chlorine ends) are negative. In each As—Cl bond, the As is δ^+ and the Cl is δ^-." The lone electron pair at the top of the pyramidal shape repels the bonding pairs. The molecule is polar.

(b)



Tetrahedral molecule: The H ends of the C—H bonds are δ^+, giving those ends partial positive charge. The nitrogen end of the C—N triple bond is δ^- and tends to have a partial negative charge, making the molecule polar.

(c)

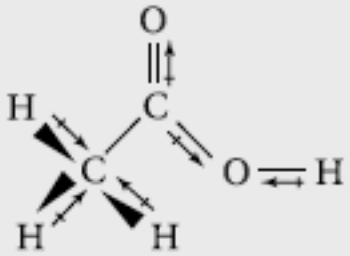


The central oxygen atom has two LPs and two BPs. The two LPs repel each other and the chlorine BPs, so the shape is bent. The chlorine ends of the bonds are δ^+ and the oxygen end is δ^-, making the molecule polar.

(d)

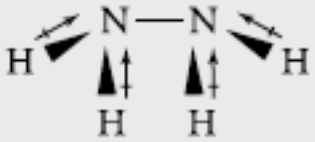
Even though each Si—Cl bond is polar, the individual Cl atoms surrounding the Si central atom exert an identical force on the bonding electrons. The Si—Cl bonds are symmetrically arranged, so the shape is tetrahedral. The polarity of the individual Si—Cl bonds neutralize each other to give a zero net polarity, so the molecule is non-polar.

(e)



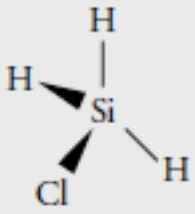
The molecule is tetrahedral around the first carbon and trigonal planar around the second carbon. The molecule is polar, with the oxygen COOH end of the molecule δ^- and the CH_3 end δ^+.

(f)



The molecule is trigonal pyramidal around each of the two nitrogen atoms. Although the formula makes it look symmetrical, the shape indicates that it is not. The molecule is polar.

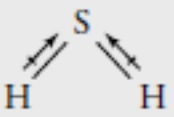
(g)



Shape: tetrahedral

Polarity: polar

(h)



Shape: bent

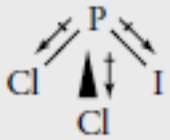
Polarity: polar

(i) $S=C=O$

Shape: linear

Polarity: polar

(j)



Shape: pyramidal

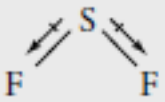
Polarity: polar

11. This molecule is based on a tetrahedral. Although the bonds themselves are slightly polar (since the difference in electronegativities between hydrogen and carbon is 0.4), the molecule itself is non-polar because it is symmetrical.

12. All polar molecules must have polar bonds, but not all non-polar molecules have non-polar bonds. If there is symmetry in a molecule such that the polarity of the individual bonds cancel each other to give a zero net charge, the molecule is non-polar. For example, CCl_4 is a non-polar molecule that has four polar bonds.

13. The number of bonding pairs and non-bonding pairs of electrons around the central atom of the molecule affects the shape of the molecule. The polarity of the individual bonds and the shape of the molecule affect its polarity.

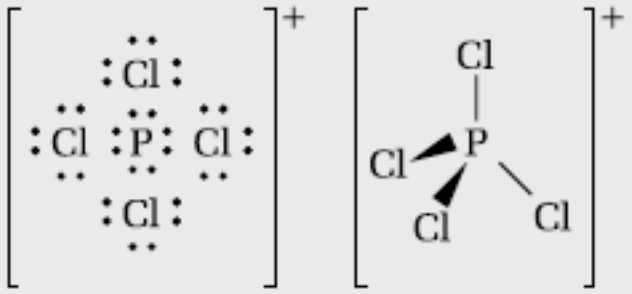
14.



Since $\Delta EN = 3.98 - 2.58 = 1.40$, the S—F bond is polar. The four electron pairs are arranged in a tetrahedral shape, and the shape of the molecule is bent. Thus, the S—F bonds do not cancel each other, and the molecule is polar.

15. Both CH_4 and CH_3OH have a tetrahedral electron pair arrangement and a tetrahedral shape. CH_4 is non-polar because of the symmetry of the molecule, with four hydrogens around the central atom. CH_3OH is polar because the polarity of the CO—OH bond is different from the polarity of the three C—H bonds.

16.



Since there is symmetry in the four P—Cl bonds, the electron pairs are arranged in a tetrahedral shape and the shape of the ion is also tetrahedral.

17. The combination reaction can be represented by the following equation:

$$BF_3 + NH_3 \rightarrow F_3BNH_3$$

Before the reaction, boron only has six electrons in the outermost energy level. The product is expected to be more stable than the reactants, because when nitrogen's lone pair is shared with boron, the covalent B—N bond completes the stable configuration of the nearest noble gas for each element in the compound.

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Q5. Intramolecular forces are the forces within molecules that hold atoms together in molecules. Intermolecular forces are the forces between molecules that hold molecules together in the liquid or solid state.

Q6. Hydrogen bonds are similar to other types of dipole-dipole interactions because they are electrostatic forces between polar molecules, like all types of dipole-dipole interactions. Hydrogen bonds are distinct from other types of dipole-dipole interactions because they are stronger and they only occur when hydrogen is bonded to nitrogen, oxygen, or fluorine, the elements with the highest electronegativities. Hydrogen bonds occur only between hydrogen nuclei and atoms of these elements on an adjacent molecule.

Q7. London (dispersion) forces account for the state at any given temperature of many substances. They are the key force of attraction between non-polar molecules. As the numbers of electrons and the size of the positive nuclear charges increase, the state of the substances at room temperature goes from gas to liquid to solid. The family of halogen (Group 17) gases is an example of this phenomenon.

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1. Dipoles are molecules with oppositely charged ends separated by short distances and are found in polar molecular structures. Dipolar molecules tend to rotate so that the positive end of one molecule is nearest to and attracts the negative end of the next nearest molecule. Dipoles tend to be attracted together.
2. Hydrogen bonding applies to molecules where a hydrogen atom is covalently bonded to a highly electronegative atom. The only combinations where this occurs are H–O, H—N, and H—F.
3. Water has only hydrogen and oxygen atoms. The rigid structure of ice keeps the molecules further apart in ice than in liquid water, where the molecules move freely and the distances between molecules change constantly. However, each polar molecule in ice is surrounded by the oppositely charged ends of other polar molecules in a hexagonal pattern. This makes the bonds between molecules stronger and more rigid and opens up a regular pattern of spaces within the hexagon, which also accounts for the increased distance between the molecules.
4. A temporary dipole is a temporary positive or negative charge distribution within the molecule. This temporary charge on one end of a molecule causes a shift in charges in an adjacent molecule. This induces a dipole in the second molecule.
5. The two factors are the number of electrons in the molecules and the shape of the molecule. The more electrons and the larger the surface area, the stronger the London (dispersion) forces.
6. Intermolecular interactions determine the distance and the direction between molecules and the strength of the bonds between molecules in substances.

7. As intermolecular forces increase, molecules are held more tightly together and the state of the substance goes from gas to liquid to solid. The melting and boiling points are a measure of the strength of these forces. The stronger the intermolecular forces, the higher the melting and boiling points.

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1. The state of a pure substance at room temperature is determined by the strength of the intermolecular forces holding it together.

2. (a) To melt an ionic compound, the ionic bonds holding the crystal lattice together must be broken.

(b) To melt a metallic solid, metallic bonds must be broken.

(c) To melt paraffin wax, a molecular solid, the intermolecular forces must be broken. Paraffin is a non-polar covalent compound, so London (dispersion) forces must be broken.

(d) Since graphite is a network solid, covalent bonds must be broken when it is melted.

3. Students will likely say that chloroform, $CHCl_3^-(l)$, is polar since not all polar bonds are equal and they therefore do not cancel. In methane, CH_4, the polarities of all bonds are equal and thus they cancel. As a result, chloroform is slightly polar and the molecules will attract each other, giving it a higher boiling point than the non-polar methane. A more significant reason is that with the more massive chlorine atoms present in chloroform, there will be more attraction between the molecules regardless of polarity; for example CCl_4, though non-polar, is still a liquid at room temperature.

4. The higher the bond energy of a substance, the higher the temperature must be to overcome the forces of attraction between the particles in the substance. The ionic and metallic bonds must be broken in order for the compounds to melt or vaporize. The energy that breaks the bonds is the kinetic energy of particles. Temperature is a measure of the average kinetic energy of atoms, ions, or molecules. Therefore, high melting points and boiling points indicate that the bonding forces are strong. In molecular substances, covalent bonds do not break when a substance melts or vaporizes. It is the intermolecular attractions that must be broken to cause melting or vaporization.

5. The relatively low melting point of 140°C suggests that the solid is a molecular solid. The solubility of the solid indicates that it is polar. Therefore, the solid is likely a polar covalent solid. To confirm, test the solution of the solid for conductivity. Low or no significant conductivity would confirm a polar covalent solid. Test also for hardness, brittleness, malleability, and ductility.

6. Ionic compounds are hard and brittle. When an object made of ionic compounds is subjected to hammering or bending, the like charges of the particles can become aligned and repel each other. This will cause the crystal structure to break along the line where the charges are repelling. Therefore, ionic compounds are not practical compounds for tools.

7. (a) The melting point of the unknown solid is likely low. Since the compound does not conduct electricity as a liquid, it is a covalent compound.

(b) As described in the answer to part (a), the compound contains covalent bonds.

8. Graphite is a solid of carbon atoms arranged as a network of hexagons layered in planes. The layers are held together by weaker bonds, so the electrons between the layers are held more loosely. When electricity is passed through the graphite parallel to the plane, these electrons in the area between the layers can move more easily, making the conductivity higher in that direction.

9. Assess the completeness and clarity of the laboratory design. Use Assesment Checklist 1.

10. Have students graph their data and mark the graph for correct labelling and following conventions. Melting and boiling points increase as the number of carbon atoms increases. This is because the boiling point and melting point of an alkane increases as the number of electrons increases. Larger molecules also have a greater surface area for Van der Waals attractions, so larger molecules have stronger attractions between them and therefore higher melting and boiling points. Smaller molecules have less surface area and fewer electrons, resulting in weaker Van der Waals interactions and lower boiling and melting points.

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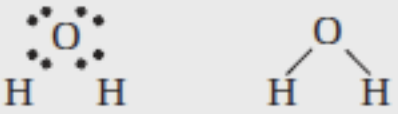
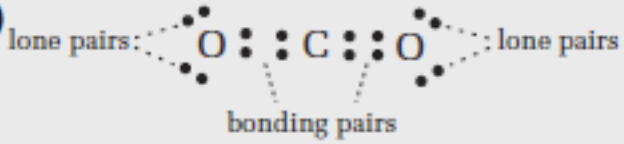
1. **(a)**



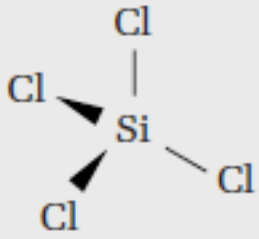
Repulsion between the lone pairs (LPs) and the bonding pairs (BPs) causes the bent structure.

(b)



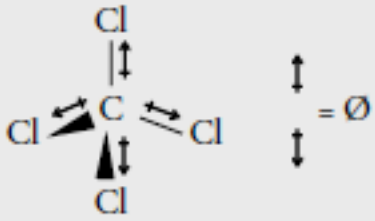
The oxygen LPs cause the oxygen atoms to be on opposite sides of the carbon atom (which has no LPs), resulting in a straight structure.

(c)



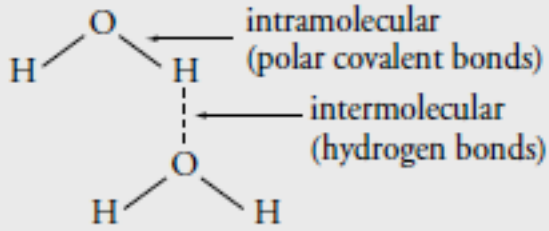
Mutual repulsion of BPs results in a tetrahedral arrangement.

(d)



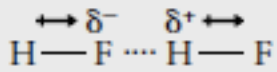
Although the bonds are polar, the polarities cancel out in this symmetrical molecule, so the molecule itself is non-polar.

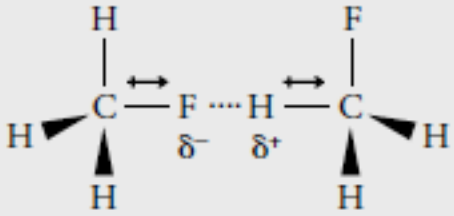
(e)



The molecules themselves are held together with polar covalent bonds (intramolecular bonds) between the hydrogen and oxygen atoms. The water molecules are held together with hydrogen bonds (intermolecular bonds).

(f)





In HF, since the hydrogen atom is bonded to the fluorine atom, hydrogen bonds will exist between molecules of hydrogen fluoride. In CH_3F, however, since the fluorine atom is bonded to the carbon and not the hydrogen atom, molecules of CH_3F will only be held together by dipole-dipole forces.

2. Oxygen, nitrogen, and chlorine are gases at room temperature because they are held together only by London (dispersion) forces. There is little attraction between the molecules because there are no dipoles; thus the molecules remain separate and the elements are gases. These elements also have relatively few electrons, compared to iodine, for example, which has more electrons and is solid at room temperature because the London (dispersion) forces are stronger.

3. Since ionic compounds do not contain molecules, it is not correct to talk about intermolecular forces when speaking of ionic substances. In ionic compounds, an ion is bound to several other ions of opposite charge in a lattice. Therefore in ionic compounds it would be correct to speak of forces between ions, or interionic forces.

4. All the halogen hydrides are polar, and atoms are held together by London (dispersion) and dipole-dipole forces. London (dispersion) forces are greatest in HI and smallest in HF. This is because the HI molecule is the largest and contains the most electrons. ΔEN is greatest in HF and smallest in HI, so the dipole-dipole forces are greatest in

HF and smallest in HI. However, ΔEN is small for halogens, so the difference in the strength of the dipole-dipole forces is not as significant as the difference in the strength of the dispersion forces. Therefore, the boiling point should increase from HF to HI. In HF the relative sizes of the hydrogen atom and the fluorine atom allow hydrogen bonding to occur. Hydrogen bonding does not occur in the other hydrogen halides. As a result, HF has the strongest intermolecular force and its boiling point is higher than the boiling points of the other hydrogen halides.

5. Polar molecules are permanent dipoles. Non-polar molecules can have induced dipoles caused by nearby ions or polar molecules that distort the electron density of the non-polar molecules. The induced dipole will disappear when the charge entity moves away or the charge disappears. A polar molecule remains polar despite changes in the electric field around it.

6. The intermolecular forces of non-polar molecules are London (dispersion) forces. The strength of these forces depends on the number of electrons and the shape and size of the molecule. Oil is a mixture of large organic molecules that have long chains of atoms and a large number of electrons. As a result, the London (dispersion) forces between the molecules are very strong, even stronger than the hydrogen bonding in water. Therefore, cooking oil has a higher boiling point than water.

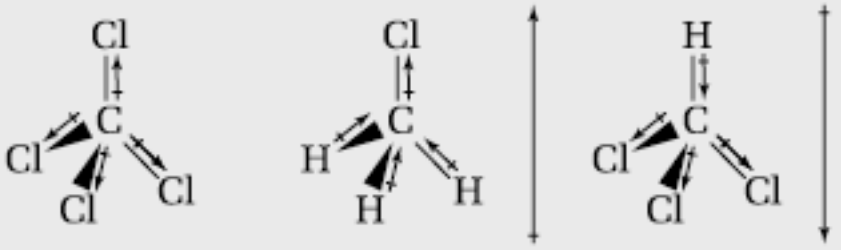
7. **(a)** London (dispersion) forces and dipole-dipole forces
   (b) ionic bonds
   (c) metallic bonds
   (d) London (dispersion) forces

8. Both $H_2O(\ell)$ and $NH_3(g)$ molecules have the same number of electrons. ΔEN is greater for O—H (1.24) than for N—H (0.84). Therefore, the O—H bond is more polar and forms stronger hydrogen bonds. Thus, hydrogen bonding is stronger in water than in ammonia.

This is also implied by the fact that water is a liquid at room temperature, whereas ammonia is a gas.

9. Order of increasing boiling points is $CH_3H_8 < C_2H_5OH < SiO_2$. SiO_2 is a network solid so it has the highest boiling point. CH_3H_8 and C_2H_5OH are covalent solids, which have lower boiling points than SiO_2. C_2H_5OH is polar, and it may have hydrogen bonding in solid form. C_3H_8 is non-polar and will not form hydrogen bonds. Therefore, C_2H_5OH probably has a higher boiling point than C_3H_8. Actual boiling-point data support this prediction: CH_3H_8 (231 K), C_2H_5OH (351 K), SiO_2 (2503 K).

10. The molecular shapes of CCl_4, CH_3Cl, and $CHCl_3$ are all tetrahedral.



CCl_4 is non-polar, but CH_3Cl and $CHCl_3$ would all have dipole-dipole forces between molecules in liquid samples. The order of the boiling points from lowest to highest is $CH_3Cl(\ell)$,) < $CHCl_3(\ell)$, and < $CCl_4(\ell)$. $CCl_4(\ell)$ has more electrons, so the London (dispersion) forces are stronger.

11. Dipole-dipole attractions are intermolecular forces between two polar molecules. An ionic bond is the attraction between two ions: one positively charged and the other negatively charged. Ionic bonds are usually thought of as intramolecular forces. In an ionic crystal, however, it is impossible to differentiate between intermolecular and intramolecular forces.

12. If the polar bonds in a molecule are arranged around the central atom in such a way that the dipoles can cancel each other, then the molecule is non-polar. Non-polar molecules with polar covalent bonds include binary (two-element) molecules with no LPs, such as CO_2, CCl_4, and BCl_3.

13. (a) PH_2F. Molecular shape is similar, so that is not a factor. Phosphorus has more electrons than nitrogen, so the London (dispersion) forces are stronger for the PH_2F. The electronegativity differences are greater between phosphorus and fluorine than they are between nitrogen and chlorine, so PH_2F is more polar and the dipole-dipole forces are greater.

(b) SiC. Silicon carbide forms a network solid, so the covalent bonds are very strong and would require very high temperatures to break. AsH_3 is a molecular substance held in the liquid state by weaker dipole-dipole and London (dispersion) forces.

(c) Xe. Xenon has more electrons than neon so its London (dispersion) forces are stronger.

(d) CH_3OH. CH_3OH has stronger hydrogen bonding than the other molecule.

(e) CH_3F. CH_3F has dipole-dipole forces but F_2 only has London (dispersion) forces.

(f) $AlCl_3$. The electronegativity differences are greater between aluminium and chlorine than they are between arsenic and chlorine, so $AlCl_3$ is more polar

and the dipole-dipole forces are greater.
Note that $AlCl_3$ sublimes at 180 °C and does not boil. It also exists in the liquid form as Al_2Cl_6 dimers.

(g) C_4H_{10}. The chlorine molecule has the same number of electrons as the C_4H_{10} molecule and both are non-polar, but the C_4H_{10} molecule has a longer shape so its London (dispersion) forces are stronger due to the larger surface area.

(h) NH_3. NH_3 has hydrogen bonds, whereas PH_3 does not.

(i) C_4H_9F. C_4H_9F is more polar so it has stronger dipole-dipole forces.

(j) Ammonia. Ammonia has hydrogen bonding.

(k) Silicon dioxide. Silicon dioxide is a network solid.

(l) Krypton has more electrons and so its London (dispersion) forces are stronger.

(m) KCl. KCl is ionic.

(n) If the formula contains a metal and a non-metal such as Mg and O, the compound is ionic. If the formula contains only non-metals such as N and H, the compound is molecular.

(o) PH_2Cl. PH_2Cl has more electrons, so the London (dispersion) forces are greater.

(p) NH_4Cl. NH_4Cl is an ionic solid.

(q) C_2H_5F. C_2H_5F is more highly polar.

(r) CH_3NH_2. CH_3NH_2 is a longer molecule, so the London (dispersion) forces are greater.

(s) C_2HCl. C_2HCl is polar, so there are dipole-dipole forces.

(t) H_2O. H_2O has strong hydrogen bonding.

(u) SiO_2. SiO_2 is a covalent solid.

(v) NaCl. NaCl is an ionic solid.

14. (a) Glycerol molecules are larger and longer than water molecules, so they would have stronger London (dispersion) forces. They also have three O—H bonds per molecule so there is a greater opportunity for hydrogen bonding. Glycerol has stronger bonds and therefore flows much more slowly.

(b) $CaCl_2$ is an ionic compound and is not an individual molecular unit.

(c) Silver forms metallic bonds and is malleable because the sea of electrons allows planes of metallic ions to slide past each other.

(d) Graphite is a network solid and so is fairly stable at high temperatures. The layers of carbon slide over each other easily, so it is an effective lubricant.

(e) In diamond, all the electrons are localized in covalent bonds and cannot move.

(f) In the molten substance, the rigid lattice structure is disrupted and the ions are free to move.

(g) The varying charges on ions result in variable melting points.

(h) The regular arrangement of carbons in a crystalline structure and the strength of the covalent bonds imply that diamond is the hardest substance.

(i) Ice has a crystalline structure that keeps the water molecules farther apart in ice than in liquid. As the water freezes, it expands, breaking the bottle.

(j) Graphite is a strong but light-weight network solid.

Answers to Applying Concepts Questions

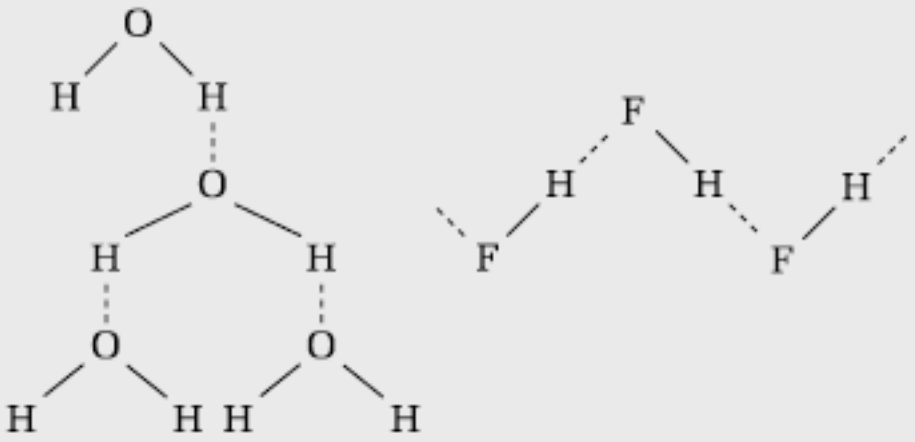
15. Measure the melting point of each substance. The ionic compound should have the higher melting point. Then test the electrical conductivity of the melted substances. The ionic material should conduct electricity, while the covalent material should not. Test the solubility of the

materials in water. If they dissolve in water, test the conductivity of the solution. The ionic material will conduct electricity in solution, while the covalent material will not.

16. Liquid A is polar since it will move towards a charged rod, indicating an opposite charge on one end of the molecule. Liquid B is non-polar since it is unaffected by a charged rod.

17. As water cools toward 4 °C, the molecules move more slowly and the attractive forces pull them closer together, so the density of the water increases. This increase in density continues until the temperature of the water is 4 °C. At 4 °C the open hexagonal structure (due to hydrogen bonding) begins to form, stopping further contraction in the volume of the water. When the water cools further, below 4 °C, the water molecules form more of the open hexagons that are found in ice. The volume begins to increase, so the density begins to decrease.

18. From electronegativity values, the H—F bond (ΔEN = 1.78) is more polar than the O—H bond (ΔEN = 1.24). This suggests that hydrogen fluoride may have a higher boiling point than water. However, water has two polar bonds whereas hydrogen fluoride has only one. The bent shape of the water molecules allows for two O—H bonds to form hydrogen bonds. Each water molecule can be linked to as many as five other molecules, but each hydrogen fluoride molecule can be linked to only two other molecules. Water boils at 100 °C and hydrogen fluoride at 19 °C. (Students would be advised to use search words "hydrogen fluoride boiling point"; using HF gives sites for the element hafnium as well.)



19. SF_2 is a bent molecule and therefore is polar because the S—F bonds are polar. SiF_4 is tetrahedral molecule. The dipoles of the tetrahedron cancel each other, making the molecule non-polar.

20. The electronegativity difference between Na and I is 1.73. NaI is an ionic solid. Therefore, the forces of attraction that must be overcome to melt this solid are electrostatic attractions due to ionic bonding. The electronegativity difference between H and I is 0.46. HI is a covalent solid that is only slightly polar. The intermolecular forces in HI are London (dispersion) forces and dipole-dipole forces which must be overcome to melt HI. These forces are much less than those for NaI.

21. (a) Diamond is a crystalline solid that does not melt. It sublimes at a temperature around 4000 K. It is one of the hardest substances known. 2,2-dimethylpropane is a liquid at room temperature. Its melting point is 257 K, and its boiling point is 283 K.

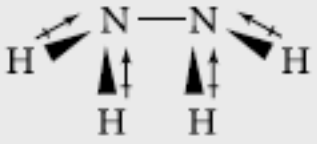
(b) The C—C bonds in both diamond and 2,2-dimethylpropane are the same. In both substances, a carbon atom is bonded to four other carbon atoms in the shape of a tetrahedron. In diamond, however, each carbon atom is bonded to four other carbon atoms in a huge network. In 2,2-dimethylpropane, each outer carbon is bonded to one carbon atom and three hydrogen atoms. Thus, 2,2-dimethylpropane exists as discrete molecules. The diamond molecule is the diamond itself. The entire crystal is a molecule because every carbon atom in the crystal is linked to four other carbon atoms within the network.

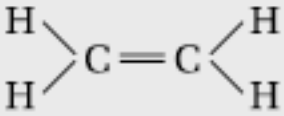
(c) The C—C bonds in diamond are networked. To cleave, to smash, or to scratch a diamond, you have to break thousands of C—C bonds. This makes diamond a very hard solid. For 2,2-dimethylpropane, the forces that hold the molecules in a solid are London (dispersion) forces. Thus, the solid is much softer and melts at a much lower temperature.

22. Carbon dioxide is a linear, non-polar covalent molecule. The only intermolecular forces that act between the molecules are London (dispersion) forces. These forces are very weak because the carbon dioxide molecule has only 22 electrons and is quite small. Therefore, carbon dioxide exists as free molecules (a gas) at room temperature. In contrast, each silicon atom is strongly bonded to four oxygen atoms in the form of a tetrahedron, and each oxygen atom is bonded to two silicon atoms to form a network. To melt silica, a lot of energy is needed to break the bonds in the network. Therefore, silica has a very high melting point (1873 K), and it exists as a solid at room temperature.

23. The ΔEN of the Se=O bond is 0.89. Thus each Se=O bond is polar. The shape of the SeO_2 molecule is bent, and a net dipole results. Therefore, the molecule is polar. In SeO_3, the selenium atom forms three double bonds with three oxygen atoms. The shape of the molecule is trigonal planar. Although the Se=O bonds are polar, the symmetry of the molecule causes the dipoles of the Se=O bonds to cancel one another. Therefore, the SeO_3 molecule is non-polar.

24. Ethene, C_2H_4, is a planar molecule with a double bond between the two carbon atoms. In this arrangement, the slight polarities of the C—H bonds cancel each other. Therefore, ethene is a non-polar molecule. The intermolecular forces in ethene are London (dispersion) forces only. Hydrazine, N_2H_4 (18 electrons), has stronger dispersion forces than ethene (14 electrons). The N—H bond is also more polar than the C—H bond. A hydrazine molecule consists of two trigonal pyramidal structures linked through the nitrogen atoms. In this arrangement, the polarities of the N—H bonds do not cancel, and the molecule is polar. As well as dipole-dipole forces and London (dispersion) forces, hydrogen bonding occurs between the hydrogen atoms of one molecule and the nitrogen atoms of another. Therefore ethene would have a lower boiling point.





25. Argon is a noble gas, and fluorine gas is a non-polar covalent molecule. Both of these substances would be held together by London (dispersion) forces. Since they have the same number of electrons, the strength of the London (dispersion) forces would be similar, so they should have similar boiling points.

26. Krypton gas atoms would be held together with London (dispersion) forces. Since London (dispersion) forces are the weakest, the boiling point of krypton should be the lowest. Hydrogen bromide gas is a slightly polar molecule, so the gas would be held together with weak dipole-dipole forces. Hydrogen selenide gas is also slightly polar and would be held together with weak dipole-dipole forces. The boiling points of HBr(g) and $H_2Se(g)$ should be similar, since they are held together with similar forces.

27. **(a)** network

(b) ionic

(c) molecular

(d) metallic

Answers to Making Connections Questions

28. There are no correct or typical answers to questions such as this. However, students should include the following ideas and provide reasons for their opinions. *Risks associated with inert substances:* Materials that are very inert present problems with disposal. A common example is foam cups, which do not decompose and therefore end up in landfill sites. If every 12-year-old bought a Kevlar™ jacket for protection, imagine the number of outgrown Kevlar™ jackets in landfill sites three years later.
Benefits associated with inert substances: Kevlar™ gloves and vests protect workers in hazardous situations. Foam cups are inexpensive and convenient for keeping drinks warm or cold.
Benefits versus risks: The benefits of Kevlar™ products outweigh the risks, since the quantity in use remains quite small. The benefits of foam cups do not outweigh the risks, since people can use re-usable or biodegradable cups instead.

29. Some examples of how intermolecular forces influence the weather: If you are standing at the top of a ski hill and snow just begins to fall, you may feel the temperature rising. As water sublimes to form snowflakes, hydrogen bonds (intermolecular forces) are formed. The making of bonds produces energy, so you may feel the air temperature rising. You get out of a swimming pool on a very hot summer day and lie on a hot concrete patio beside the pool. You find yourself shivering. As the water on your body evaporates, the hydrogen bonds in the water need energy in order to break, and absorb heat from your body, making you cooler.

On a cold winter morning, you may notice frost on classroom windows. During the previous night, the glass window became very cold. As water vapour passed over the window, it cooled and the water molecules were brought closer together. Hydrogen bonds formed to make small crystals of ice. The frost "painting" on the window was formed by water vapour subliming to the solid. Because of the tremendous amount of heat necessary to change the state of water, it has a moderating effect on the temperatures. The average of all winter and summer temperatures for Edmonton is 2 °C, for Fort Chipewyan it is –2 °C, for Calgary it is 3 °C, and for Medicine Hat it is 5 °C. Vancouver has the highest average in Canada, around 10 °C.

30. (a) Although atoms do not become charged as do clouds, a separation of charge may form in an atom due to electron movement. As with a cloud, the negative "part" of an atom will lead to a separation of charge in nearby atoms, just like the negatively charged cloud causes a positive leader. The opposite charges in

the atoms lead to the attractive London (dispersion) force. Similarly, the opposite charges in the cloud and in the positive leader result in the leader moving (being attracted) to the cloud.

(b) The separation of charge in an atom leads to a separation of charge in nearby atoms and the resulting London (dispersion) force. This is similar to what happens when people are near a charged cloud. The charged cloud leads to a separation of charge in the nearby peoples' hair, causing an attraction between the hair and the cloud, and causing their hair to stand on end.